

## FIFTY FREQUENTLY FORGOTTEN FUN FACTS

This packet contains topics from each of the units we worked on this year with questions. Most of the questions are similar to what you would expect to see on Part B2 and C of the Regents Exam in Chemistry. The multiple choice questions mirror common questions found on Parts A and B1.

### I. ATOMIC STRUCTURE & NUCLEAR CHEMISTRY

1) Protons are +1 each with a mass of 1 amu each, the number of protons = atomic number, nuclear charge = + (# protons). [Periodic Table]

- a) How many protons are there in a nucleus of Kr-85? 36
- b) What is the nuclear charge of an atom of Br? +35
- c) What is the mass of the protons in a nucleus of O-15? 15 9

2) Neutrons are neutral with a mass of 1 amu each, # neutrons = mass # - atomic number. Isotopes = atoms of the same element (same atomic #) but different # of neutrons (mass #). [Periodic Table]

- a) How many neutrons are there in the nucleus of  $^{56}_{26}\text{Fe}$ ? 30 # neutrons = ATOMIC MASS - ATOMIC NUMBER
- b) Circle the two nuclei that are isotopes of each other:  $^{15}_8\text{O}$   $^{15}_7\text{N}$   $^{16}_8\text{O}$   $^{16}_9\text{F}$   
 ↳ same atomic #, different atomic mass.

3) Electrons are each -1 with a mass that is VERY, VERY tiny compared to the mass of a proton or neutron.

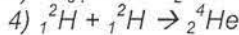
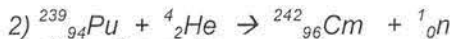
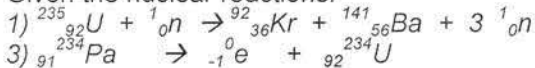
- (basically zero)
- a) Which particle has a mass that is  $1/1836^{\text{th}}$  the mass of a proton?
- 1)  $^4_2\text{He}$       2)  $^1_1\text{H}$       3)  $^0_{-1}\text{e}$       4)  $^1_0\text{n}$

4) Natural Decay: Parent Nuclide → Decay particle + daughter nuclide [Tables N and O]

- a) Write the decay for U-238:  $^{238}_{92}\text{U} \rightarrow ^4_2\text{He} + ^{234}_{90}\text{Th}$
- b) Write the decay for K-37:  $^{37}_{19}\text{K} \rightarrow ^0_{-1}\text{e} + ^{37}_{18}\text{Ar}$
- c) Write the decay for P-32:  $^{32}_{15}\text{P} \rightarrow ^0_{-1}\text{e} + ^{32}_{16}\text{S}$

5) Artificial Transmutation is when a relatively stable nucleus is impacted by a particle bullet at high speeds and becomes an unstable nucleus of a different element. Nuclear fission occurs when nuclei of U-235 or Pu-239 are impacted by a neutron and split into two smaller nuclei and more neutrons. Nuclear fusion occurs when two small nuclei of hydrogen combine at high temperatures and pressures to form larger nuclei of helium. Both fission and fusion convert mass into a huge amount of energy.

Given the nuclear reactions:



- a) Which reaction represents natural decay? 3
- b) Which reaction represents artificial transmutation? 2 (FUSION - NATURAL SUN)
- c) Which reaction represents nuclear fission? 1
- d) Which reaction represents nuclear fusion? 4

## NOT ON REFERENCE TABLE

6) Weight-average mass =  $\frac{(\% \text{ of isotope 1} \times \text{mass of isotope 1})}{100} + \frac{(\% \text{ of isotope 2} \times \text{mass of isotope 2})}{100} + \dots$

a) What is the weight-average mass of an isotope if X-50 (mass = 50.0 amu) has an abundance of 20.0% and X-52 (mass = 52.0 amu) has an abundance of 80.0%? Show all work:

$$\frac{(20 \times 50)}{100} + \frac{(80 \times 52)}{100} =$$

Answer must be between the highest and lowest

$$10 + 41.6 = 51.6 \text{ g}$$

answer: 51.6 amu

7) # Half-lives = (time elapsed / length of half-life) [Tables N and T]

a) A sample of Co-60 is left to sit for 15.78 years. How many half-lives have gone by?

$$\frac{\text{Total Time}}{\text{Half-life}} = \frac{15.78}{5.26} = \boxed{3 \text{ half-lives}}$$

b) What percent of the original sample remains after this number of half-lives?

$$1 - \frac{1}{2} - \frac{1}{4} - \frac{1}{8}$$

Remember - Count jumps!

12.5%

c) If the original mass of the sample was 20.0 grams, how many grams of Co-60 remain?

$$20.0 \text{ g} \rightarrow 10.0 \text{ g} - 5.0 \text{ g} - \boxed{2.5 \text{ g}}$$

## II. PHYSICAL BEHAVIOR OF MATTER

Fuse - melt

8) Heat of Fusion = heat added to MELT or heat removed to FREEZE a substance.  $q = m H_f$  [Tables B, T]

a) How many joules are required to melt 10.0 grams of water at the melting point? Show all work:

$$q = m H_f \quad (10.0 \text{ g})(334 \text{ J/g}) = 3340 \text{ J}$$

9) Heat of Vaporization = heat added to BOIL or removed to CONDENSE a substance.  $q = m H_v$  [Tables B, T]

a) How many joules are required to boil 20.0 grams of water at the boiling point? Show all work:

$$q = m H_v \quad (20.0 \text{ g})(2260 \text{ J/g}) = 45,200 \text{ Joules}$$

10) Calorimetry:  $q = mc\Delta T$  = heat that is added or removed to change the temperature of a substance, but NOT its phase. [Tables B, T]

a) How many joules are required to raise the temperature of 15.0 grams of water from 10.0°C to 25.0°C? Show all work:

$$q = m c \Delta T \quad (15.0 \text{ g})(4.18 \text{ J/g}^\circ\text{C})(25.0 - 10.0^\circ\text{C}) = 940.5 \text{ Joules}$$

b) 50.0 grams of water absorb 1000. J of energy. By how much does the temperature increase? Show all work:

$$q = m c \Delta T$$

$$1000 \text{ J} = (50.0 \text{ g})(4.18 \text{ J/g}^\circ\text{C})(\Delta T)$$

$$\Delta T = \frac{1000}{(50.0 \times 4.18)} = 4.8^\circ\text{C}$$

**11) Gas Laws: Temperature must be in Kelvin, STP is found on Reference Table A. [Tables A, T]**

a) 50.0 mL of a gas at STP is heated to 400.0°C and is compressed to 20.0 mL. What is the new pressure of the gas?  
Show all work:

$$\frac{P_1 V_1}{T_1} = \frac{P_2 V_2}{T_2}$$

$$P_1 = 1 \text{ atm}$$

$$V_1 = 50.0 \text{ mL}$$

$$T_1 = 273 \text{ K}$$

$$P_2 = x$$

$$V_2 = 20.0 \text{ mL}$$

$$T_2 = 400 + 273 = 673 \text{ K}$$

$$P_2 = \frac{P_1 V_1 T_2}{T_1 V_2}$$

$$= \frac{(1 \text{ atm})(50.0 \text{ mL})(673 \text{ K})}{(273 \text{ K})(20.0 \text{ mL})} = 6.2 \text{ atm}$$

**12) Avogadro's Hypothesis -- When ANY two gases are at the same T and P, they will have the same volume and THEREFORE the same number of molecules.**

a) Which of the following samples of gas contain the same number of molecules?

Gas	Pressure	Temperature	Volume
A	100 kPa	300. K	50.0 mL
B	100 kPa	300. K	50.0 mL
C	200 kPa	200. K	100.0 mL
D	200 kPa	200. K	50.0 mL

Answer: A and B

**13) Temperature (a measure of the KE) remains constant during a phase change, only PE changes during a phase change (Heat of Fusion or Vaporization).**

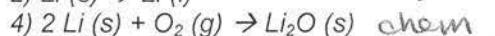
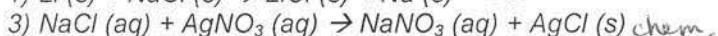
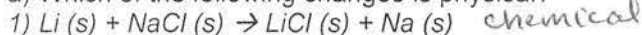
Given the following data table:

Time (min)	0	1	2	3	4	5	6	7	8	9	10	11	12	13	14	15	16	17	18
Temp (°C)	70	75	80	80	80	80	89	98	107	116	116	116	116	116	116	136	156	186	206

- s → s → l → g
- a) What is the melting point of this substance? 80°C
- b) What is the boiling point of this substance? 116°C
- c) Between minute 0 and 2, what is happening to kinetic energy? KE ↑
- d) Between minute 9 and 14, what is happening to kinetic energy? KE — (same)
- e) Between minute 5 and 9, what is happening to potential energy? PE — same
- f) Between minute 2 and 5, what is happening to potential energy? PE ↑

**14) Phase changes and dissolving are physical changes.**

a) Which of the following changes is physical?



### III. PERIODIC TABLE AND BONDING

15) Elements Br, I, N, Cl, H, O and F form diatomic molecules through nonpolar covalent bonding when there are no other elements present.

a) Complete the following reaction:  $2 \text{Na} + 2 \text{HOH} \rightarrow 2 \text{NaOH} + \text{H}_2$

b) Complete the following reaction:  $2 \text{FeCl}_3 \rightarrow 2 \text{Fe} + 3 \text{Cl}_2$

16) Noble gases are nonreactive, forming monatomic molecules. [Periodic Table]

a) Name an element that exists as monatomic molecules: Ne

17) When metal atoms form ions, they lose all their valence electrons, and their dot diagrams are the metal symbol, in brackets, with no dots and the + charge on the upper right, outside the brackets. [P.T.]

a) What is the electron configuration of a  $\text{K}^{+1}$  ion? 2-8-8

b) A  $\text{Ca}^{+2}$  ion has the same electron configuration as which noble gas? Ar

c) When Fe forms a +2 ion, its radius decreases

d) Draw the dot diagram for the  $\text{Li}^{+1}$  ion:



18) When nonmetal atoms form ions, they gain enough electrons to have a stable octet (8 valence electrons), and their dot diagrams are the nonmetal symbol, in brackets, with 8 dots and the - charge on the upper right, outside the brackets. [Periodic Table]

a) What is the electron configuration of a  $\text{Cl}^{-1}$  ion? 2-8-8

b) A  $\text{S}^{-2}$  ion has the same electron configuration as which noble gas? Ar

c) When O forms a -2 ion, its radius increases

d) Draw the dot diagram for the  $\text{F}^{-1}$  ion:



19) Hydrogen bonds are strongest between molecules with the greatest electronegativity difference. [Table S]

a) Which molecule has the strongest hydrogen bond attractions? 1) HF 2) HBr 3) HCl 4)  $\text{H}_2\text{O}$

20) Ionic character increases as electronegativity difference increases. [Table S]

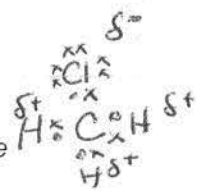
a) Which compound has the greatest ionic character? a) NaBr b) NaI c) NaCl d) NaF

biggest difference  
in electronegativity

21) At STP, the liquids on the Periodic Table are Br and Hg. The gases are N, Cl, H, O, F and the Noble Gases. All other elements are solids. [Periodic Table]

- a) Which element on the Periodic Table is a nonmetallic liquid at STP? Br<sub>2</sub>
- b) Which element at STP is a liquid that conducts electricity well? Hg
- c) Name an element that exists in a crystal lattice at STP: Na (BAD question)
- d) Name an element that has no definite volume or shape at STP: O<sub>2</sub> (gas)

22) Electronegativity is an atom's attraction to electrons in a chemical bond. [Table S]

- a) Which element, when bonded with O, will form the partially negative end of a polar covalent bond? H H<sub>2</sub>O
- b) Which element has the greatest attraction to electrons when bonded to Na?
- 1) N 3.0      2) O 3.5      3) S 2.6      4) Al 1.6
- c) In the molecule CH<sub>3</sub>Cl, which element represents the partially negative end of the molecule?
- 1) C      2) H      3) Cl      4) none, it's a nonpolar molecule
- 

23) Ionization energy is the energy required to remove the most loosely held valence electron from an atom in the gas phase. [Table S]

- a) Four elements are heated at the same rate. Which will lose an electron first?
- 1) Na 496      2) Br 1140      3) Fe 762      4) Ca 590

24) Polyatomic ions form ionic bonds with other ions, but are themselves held together by covalent bonds. [Table E]

- a) Which of the following compounds contains both ionic and covalent bonds?
- 1) NaCl      2) CH<sub>4</sub>      3) CaCO<sub>3</sub>      4) CO<sub>2</sub>
- ionic      covalent      ionic + covalent      covalent

## IV. COMPOUNDS

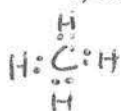
25) Ionic compounds are made of a metal and nonmetal, or a metal and a negative polyatomic ion. They have high melting points, and conduct electricity when dissolved in water (electrolytes) or melted. [P. T.]

- a) Which of the following substances is the best conductor of electricity when dissolved in water? (polar)
- 1) K<sub>2</sub>SO<sub>4</sub>      2) CCl<sub>4</sub>      3) C<sub>6</sub>H<sub>12</sub>O<sub>6</sub>      4) NO<sub>2</sub>
- ionic      non-polar      non-polar      non-polar

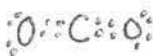
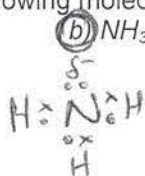
26) Molecular compounds tend to be soft, have low melting points and high vapor pressures. Hydrogen bonds are the strongest of the intermolecular forces (when the H of one polar molecule attracts the N, O or F of another polar molecule), followed by dipole (where the more electronegative end of one polar molecule attracts the less electronegative end of another polar molecule) and London Dispersion forces are the weakest, where motion of electrons through the molecule causes temporary poles to form. Molecular substances (with the exception of acids) are poor conductors of electricity (nonelectrolytes). [P. T.]

- a) Which of the following substances is the poorest conductor of electricity when dissolved in water?
- 1) CaCl<sub>2</sub>      2) HCl      3) NO<sub>2</sub>      4) NaBr

- b) Which of the following molecules is subject to hydrogen bond attractions in the solid and liquid phase?
- 1) CH<sub>4</sub>      2) NH<sub>3</sub>      3) CO<sub>2</sub>      4) C<sub>3</sub>H<sub>8</sub>



5



27) Network solids are substances that do not have distinct molecules or ions that can separate with heating. To melt a network solid, covalent bonds have to be broken. This takes tremendous energy, meaning that network solids have extremely high melting points. They are insoluble in water, and are poor conductors of electricity. Examples of network solids are diamond (C), sapphire, ruby, corundum ( $\text{Al}_2\text{O}_3$ ) and quartz ( $\text{SiO}_2$ ).

a) Which of the following is a network solid?

1)  $\text{NaCl}$

b)  $\text{H}_2\text{O}$

c)  $\text{SiO}_2$

d)  $\text{Hg}$

28) ONLY metals with more than one listed charge need a Roman numeral after their name (Stock system) when naming an ionic compound. Nonmetals with more than one oxidation state will also need a Roman numeral in their name if they are the less electronegative atom in a molecular compound. [P. T., Table E]

a) Name the compound  $\text{Cu}(\text{NO}_3)_2$ : Copper (II) Nitrate

b) Write the formula for iron (III) sulfite:  $\text{Fe}^{+3} \text{SO}_3^{-2} \text{Fe}_2(\text{SO}_3)_3$

c) Name the compound  $\text{NO}_2$ , using the Stock system: Nitrogen (IV) oxide

d) Write the formula for phosphorous (IV) oxide:  $\text{P}^{+4} \text{O}^{-2} \text{P}_2\text{O}_4$   ~~$\text{PO}_2$~~  <sup>EMPIRICAL FORMULA</sup>

29) Formula Mass = sum of all atomic masses in the compound, rounded to the tenths place, with the units g/mole. [Periodic Table]

a) Determine the formula mass of  $\text{Cu}(\text{NO}_3)_2$ :  $\text{Cu} \quad 2 \times \text{N} \quad 6 \times \text{O}$   
 $63.546 + (2 \times 14.0067) + (6 \times 15.994) = 187.5234 \text{ g/mole}$

30) grams / formula mass = moles      moles X formula mass = grams [Periodic Table, Table T]

a) Using the formula mass of  $\text{Cu}(\text{NO}_3)_2$ , how many moles are there in 100.0 grams of  $\text{Cu}(\text{NO}_3)_2$  (show all work):

$$\# \text{ moles} = \frac{\text{given}}{\text{gfm}} = \frac{100 \text{ g}}{187.5234 \text{ g}} = .53 \text{ moles}$$

b) Using the formula mass of  $\text{Cu}(\text{NO}_3)_2$ , how many grams are there in 2.5 moles of  $\text{Cu}(\text{NO}_3)_2$  (show all work):

$$\# \text{ moles} = \frac{\text{given}}{\text{gfm}} \quad 2.5 \text{ moles} = \frac{x}{187.5234 \text{ g/mole}} \quad x = 468.8 \text{ g}$$

31) Molecular Formula = (Molecular Mass / Empirical Mass) X Empirical Formula [Periodic Table]

a) Quantitative analysis determines that a compound has an empirical formula of CH and a molecular mass of 26 grams/mole. Determine the molecular formula of this compound, showing all work:

$$\text{CH} = 12.0111 + 1.00794 = 13$$

$$\frac{\text{MM}}{\text{EM}} = \frac{26}{13} = 2 \quad (\text{CH})_2 = \text{C}_2\text{H}_2$$

32) % Of Water In A Hydrate = (mass of water / mass of hydrate) X100 [Periodic Table, Table T]

a) What is the % by mass of H<sub>2</sub>O in CaCl<sub>2</sub> • 2 H<sub>2</sub>O? Show all work:

$$\begin{array}{l} \text{Ca } 1 \times 40.08 = 40.08 \\ \text{Cl } 2 \times 35.453 = 70.906 \\ \text{H}_2\text{O } 2 \times 18.0 = 36.0 \\ \hline 146.986 \end{array}$$

$$\% = \frac{\text{mass of part}}{\text{mass of whole}} \times 100 = \frac{36.0}{146.986} \times 100 = 24.49\% \text{ H}_2\text{O}$$

b) 2.00 grams of hydrate are heated to a constant mass of 1.20 grams. What was the % by mass of water in the hydrate? Show all work:

$$\% = \frac{\text{mass of part}}{\text{mass of whole}} \times 100 = \frac{0.8 \text{ g}}{2.0 \text{ g}} \times 100 = 40\% \text{ H}_2\text{O}$$

$$2.00 - 1.20 = 0.8 \text{ g H}_2\text{O}$$

## V. REACTIONS

33) Synthesis, Decomposition, and Single Replacement reactions are all examples of REDOX reactions, because one species is oxidized and another is reduced. Double replacement (including neutralization) reactions are NOT redox reactions.

a) Which of the following reactions is an example of a redox reaction?

- 1)  $\text{NaCl (s)} \rightarrow \text{Na}^+ \text{ (aq)} + \text{Cl}^- \text{ (aq)}$       2)  $2 \text{ K (s)} + \text{CaSO}_4 \text{ (aq)} \rightarrow \text{K}_2\text{SO}_4 \text{ (aq)} + \text{Ca (s)}$   
 3)  $\text{Ca(NO}_3)_2 \text{ (aq)} + \text{K}_2\text{CO}_3 \text{ (aq)} \rightarrow \text{CaCO}_3 \text{ (s)} + 2 \text{ KNO}_3 \text{ (aq)}$       4)  $\text{H}_2\text{O (l)} \rightarrow \text{H}_2\text{O (g)}$

34) The driving force behind double replacement reactions is the formation of an insoluble precipitate as one of the products. [Table F]

a) Is PbCl<sub>2</sub> soluble or insoluble? Explain, based on Table F:

halides are soluble except when they combine with  $\text{Pb}^{+2}$  and is insoluble

b) In the reaction  $\text{Li}_2\text{SO}_4 + \text{Ba(NO}_3)_2 \rightarrow \text{BaSO}_4 + 2 \text{ LiNO}_3$ , write the formula for the precipitate: BaSO<sub>4</sub>

35) Stoichiometry: moles of given X (coeff. of target / coeff. of given) = moles of target

a) For the reaction  $\text{CH}_4 + 2 \text{ O}_2 \rightarrow \text{CO}_2 + 2 \text{ H}_2\text{O}$ , how many moles of H<sub>2</sub>O are formed when 20.0 moles of CH<sub>4</sub> are burned? Show all work.

$$\begin{array}{l} \text{CH}_4 : \text{H}_2\text{O} \\ 1 : 2 \\ 20 : x \end{array}$$

$$\frac{1}{20} = \frac{2}{x} \quad x = 40.0 \text{ moles}$$

## VI. KINETICS & EQUILIBRIUM

36) Energy is absorbed to break chemical bonds and released when new bonds are formed.

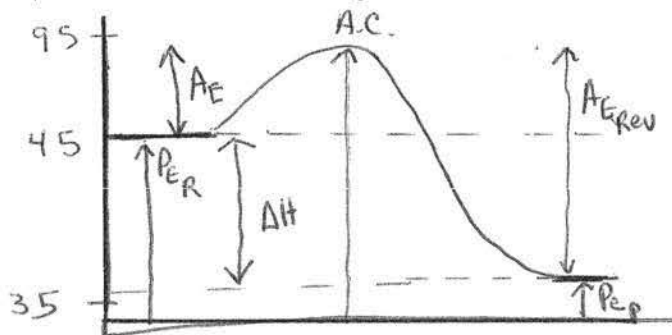
a) Which statement best describes the reaction  $H + H \rightarrow H_2 + \text{energy}$ :

- 1) A bond is being broken, which absorbs energy      2) A bond is being formed, which absorbs energy  
3) A bond is being broken, which releases energy      4) A bond is being formed, which releases energy

37) Activation energy is the energy given to the reactants to get the reaction started.

If the heat of reactants are 45 KJ, the heat of the products are 35 KJ and the heat of the activated complex is 95 KJ,

- a) What is the activation energy of this reaction? 50 kJ
- b) Adding a catalyst will lower the activation energy by lowering steps from the reaction pathway (mechanism).
- c) Adding an inhibitor will raise the activation energy by raising steps to the reaction pathway.
- d) The heat of reaction ( $\Delta H$ ) of this reaction is  $H_p - H_R$   $35 - 45 = -10 \text{ kJ}$
- e) Sketch and label a PE diagram for this reaction:



$$\Delta H = H_p - H_R$$

38) At equilibrium, the RATES are equal. The amounts don't have to be.

a) For the change  $H_2O(l) + \text{heat} \rightleftharpoons H_2O(g)$  at  $100^\circ\text{C}$ , what must be true about the rate of boiling and the rate of condensing?

They must equal rates

39) In Le Chatelier's Principle, if a system is at equilibrium, if something is added, then the equilibrium will shift away from the side it is on. If something is removed, then the equilibrium will shift towards that side. After the shift, whatever is being shifted towards will increase in concentration, and whatever is being shifted away from will decrease in concentration.

For the equilibrium  $N_2(g) + 3H_2(g) \rightleftharpoons 2NH_3(g) + \text{heat}$ :

- a) If  $N_2$  is added, which way will the equilibrium shift? right  $\rightarrow$
- b) If temperature is decreased, which way will the equilibrium shift? left  $\leftarrow$
- c) If pressure is increased, which way will the equilibrium shift? right  $\rightarrow$  (towards less)
- d) If  $H_2$  is removed, what will happen to the concentration of  $NH_3$ ? decrease  $\downarrow$
- e) If  $NH_3$  is added, what will happen to the concentration of  $N_2$ ? increase  $\uparrow$

## VII. SOLUTIONS

40) Solubility is a measure of how many grams of solute are required to saturate a given amount of solvent at a given temperature. [Table G]

a) How many grams of  $\text{NH}_4\text{Cl}$  are required to saturate a 100-gram sample of water at  $30^\circ\text{C}$ ? 42 g

b) What is the solubility of  $\text{KNO}_3$  in 50.0 grams of water at  $60^\circ\text{C}$ ? 53g (106 per 100 mL  $\text{H}_2\text{O}$ )

41) Molarity = moles / L, if grams are given, convert to moles, if mL are given, convert to L. [Table T]

a) What is the molarity of a solution of NaOH (formula mass = 40.0 g/mole) if it contains 20.0 grams of NaOH dissolved into 400.0 mL of solution? Show all work:

$$M = \frac{\text{moles}}{\text{Liters}}$$

$$\# \text{ moles} = \frac{\text{given}}{\text{gfm}}$$

$$M = \frac{0.5 \text{ moles}}{0.4 \text{ L}} = 1.25 \frac{\text{moles}}{\text{L}} \text{ or } M$$

$$\text{H moles} = \frac{20}{40} = .5 \text{ moles}$$

42) moles = Molarity X L. If asked for grams, convert moles to grams at the end. [Table T]

a) How many grams of NaOH (formula mass = 40.0 g/mole) are needed to make 500.0 mL of a 0.200 M solution of NaOH? Show all work:

$$\text{moles} = (.200 \text{ M})(.500 \text{ mL}) = .1 \text{ moles}$$

$$\# \text{ mole} = \frac{\text{given}}{\text{g fm}}$$

$$\bullet \text{ Moles} = \frac{x}{40.0 \text{ g/mol}}$$

4 grams

43) When a solute is dissolved in water, the boiling point of the solution increases and the freezing point of the solution decreases as the concentration increases. The more ions the solute creates upon dissolving the greater the increase in boiling point/decrease in freezing point. Electrolytes (ionic compounds and acids) put ions into solutions, nonelectrolytes (molecular substances) don't.

a) Which solution of NaCl (aq) has the highest boiling point? 1) 1.0 M 2) 2.0 M 3) 3.0 M 4) 4.0 M

b) Which 1.0 M solution has the lowest freezing point? 1) NaCl 2) CH<sub>4</sub> 3) CaCO<sub>3</sub> 4) MgCl<sub>2</sub>

### III. ACIDS AND BASES

44) Use  $M_a V_a = M_b V_b$  ONLY for titration problems, where they give information on BOTH the acid and base. If it is not a titration problem, and they ask for the molarity, use  $\text{Molarity} = \text{moles} / \text{L}$ . [Table T]

a) 50.0 mL of 3.0 M HCl are required to neutralize 30.0 mL of an NaOH solution. What is the molarity of the NaOH? Show all work:

$$M_A V_A = M_B V_B$$

$$(3.0M)(50.0mL) = (M_B)(30.0mL)$$

$$M_B = \frac{(3.0)(50.0)}{(30.0)} = 5M$$

b) A solution of NaOH contains 2.0 moles dissolved into 4.0 L of solution. What is the molarity of the NaOH solution?  
Show all work:

$$\text{Molarity} = \frac{\text{moles}}{\text{liters}} = \frac{2.0 \text{ moles}}{4.0 \text{ L}} = 0.5 \text{ M}$$

45) Bronsted/Lowry Acids are proton donors (give off  $H^+$ ) and B/L Bases are proton acceptors (pick up  $H^+$ ).

a) In the reaction  $NH_3 + HCl \rightleftharpoons NH_4^+ + Cl^-$ , the B/L acid in the forward reaction is: HCl

b) In the reaction  $HCl + H_2O \rightleftharpoons H_3O^+ + Cl^-$ , the B/L base in the reverse reaction is:  $Cl^-$

## IX. ELECTROCHEMISTRY

46) ALL species identified in a redox reaction MUST have their charges written. Be sure to indicate whether the charge is positive (+) or negative (-), as well as the numeric value of the charge. [P. T., Table E]

a) For the reaction  $2 Na^0 + 2 HCl^{+1-1} \rightarrow 2 NaCl^{+1-1} + H_2^0$  LEO GER

Write the charges of each species above their symbols in the above reaction

Oxidation half-reaction:  $2 Na^0 \rightarrow 2 Na^{+1} + 2 e^-$

Reduction half-reaction:  $2 H^{+1} + 2 e^- \rightarrow H_2^0$

Oxidizing Agent:  $H^{+1}$  Reducing Agent:  $Na^0$

Spectator Ion:  $Cl^{-1}$

b) What is the negative ion found in a solution of nitric acid?  $NO_3^{-1}$   
 $HNO_3$

47) The sum of all the charges of each element in a compound is zero. Oxygen is always -2 (unless it is part of the peroxide ion,  $O_2^{2-}$ , in which case O is -1). Any element by itself has a charge of 0. [P. T., Table E]

a) What is the charge of Cl in  $CaCl_2$ ?  $Ca^{+2} Cl^{-1}_2$

b) What is the charge of Cl in  $Cl_2$ ?  $\emptyset$  diatomic

c) What is the charge of Cl in  $Ca(ClO_2)_2$ ?  $Ca^{+2} (Cl^{+3} O_2^{-2})^{-1}$  (+3)

48) Voltaic cells produce electricity using a spontaneous redox reaction, electrolytic cells use electricity to decompose compounds containing Group 1, 2 or 17 elements. [Table J, P. T.]

a) A voltaic cell has Al and Au as its metal electrodes. Which metal acts as the anode? Al (more active)

b) A voltaic cell has Fe and Sn as its metal electrodes. From which metal to which metal will electrons flow?

From Fe to Sn. (Anode to Cathode)

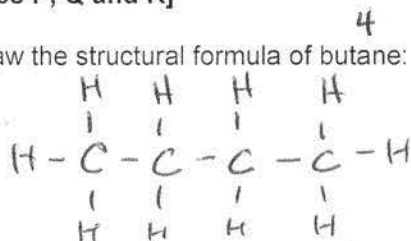
c) Name a metal that can be formed by electrolytic reduction. Group 1 or 2

d) Name a nonmetal that can be formed by electrolytic oxidation.  $Cl_2$   $O_2$   $H_2$  (Halogens)

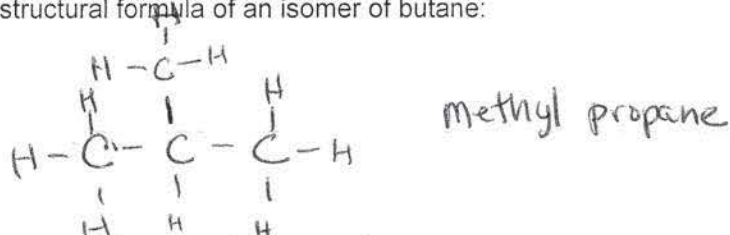
## X. ORGANIC CHEMISTRY

49) Isomers are organic compounds with the same molecular formula, but with a different structural formula. [Tables P, Q and R]

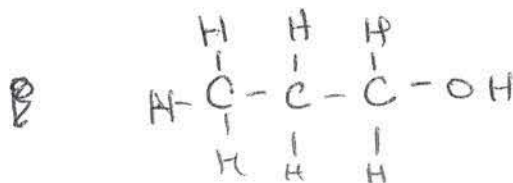
a) Draw the structural formula of butane:



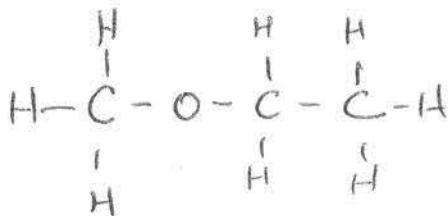
b) Draw the structural formula of an isomer of butane:



c) Draw the structural formula of 1-propanol:



d) Draw the structural formula of an ether that is an isomer of 1-propanol:



50) Addition reactions involve alkenes or alkynes. Substitution reactions involve alkanes. Use Reference Table Q to determine which type of hydrocarbon you have. [Table Q]

a) Which of the following molecules can undergo a addition reaction?

1)  $\text{C}_3\text{H}_8$

alkane

2)  $\text{C}_4\text{H}_8$

alkene

3)  $\text{C}_5\text{H}_{12}$

alkane

4)  $\text{CH}_4$

alkane

# ANSWER KEY

## Regents Exam In Chemistry Review Homework #1

Name \_\_\_\_\_

1) How many protons are in an atom of iron? 26 protons

2) How many neutrons are in the nucleus of Ca-41? 21 neutrons  ${}^{41}_{20}\text{Ca}$

3) What is the most common isotope of argon? Argon-40

4) What is the nuclear charge of an atom of calcium? +20

5) What is the mass of an electron?  $\frac{1}{1836}$  or  $\approx 0$

6) Based on Reference Table N, write the decay equation for Tc-99  ${}^{99}_{43}\text{Tc} \rightarrow {}^0_{-1}\text{e} + {}^{99}_{44}\text{Ru}$

7) What is nuclear fusion? A nuclear reaction in which two light nuclei combine to form a heavier nuclei  ${}^1_1\text{H} + {}^1_1\text{H} \rightarrow {}^2_2\text{He}$

8) Nuclear reactions give off thousands of times more energy than chemical reactions. Where does this energy come from?

Mass is converted into energy

9) Draw the dot diagram for an atom of N:  $\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{N}}}\cdot$

10) Draw the dot diagram for an ion of  $\text{N}^{3-}$ :  $\cdot\overset{\cdot\cdot}{\underset{\cdot\cdot}{\text{N}}}^{\cdot\cdot\cdot}$

11) Explain how a photon of light is formed: An electron absorbs energy and jumps up to excited state and when it jumps back down to a lower energy level, it releases a photon - (packet of light)

12) Ernest Rutherford shot alpha particles at gold foil. What happened to the alpha particles?

Most passed right through, a few bounced back.

13) What did this show about the structure of the atom?

The atom is mostly empty space, nucleus is dense and positive

14) Write an electron configuration for oxygen that is in the excited state. 1-6-1

15) An atomic mass unit (amu) is defined as what fraction of what isotope's mass?  $\frac{1}{12}$  mass of a carbon atom

## Regents Exam In Chemistry Review Homework #2

Name \_\_\_\_\_

1) What is the geometric shape that solid substances are found in called? crystals

2) Why do ionic liquids conduct electricity, while ionic solids do not? The ions are not mobile in a solid. In a solution, they are mobile and able to conduct electricity

3) Two samples of different gases each occupy 4.0 L at STP. What is true about the number of molecules contained in each of the two samples?

Equal volumes of gases have equal number of molecules

4) What is the vapor pressure of ethanol at a temperature of 50°C? 30 kPa "H"

5) What is the boiling point of propanone under a pressure of 20 kPa? 16°C "H"

6) What is the normal boiling point of ethanoic acid? STP H 117°C - - - -

7) What happens to the boiling point of water if a solute is dissolved into it? The boiling point increases

8) What happens to the melting point of water if a solute is dissolved into it? The melting point decreases

9) As temperature increases, pressure on a sample of confined gas will increase

"S" 10) Give two examples of physical properties: boiling point and density

11) Give two examples of chemical changes: burning a match and formation of a gas

12) Why do metals conduct electricity? They have a mobile sea of electrons

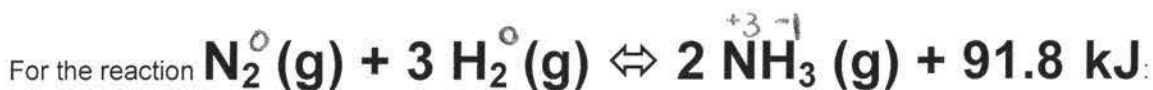
13) What three types of substances are able to conduct electricity? metals, ionic substances dissolved in water, electrolyte.

14) How many valence electrons do all ions (except H, Li, Be and B) have? 8

15) This many valence electrons is called full outer shell. (OCTET)

## Regents Exam In Chemistry Review Homework #3

Name \_\_\_\_\_



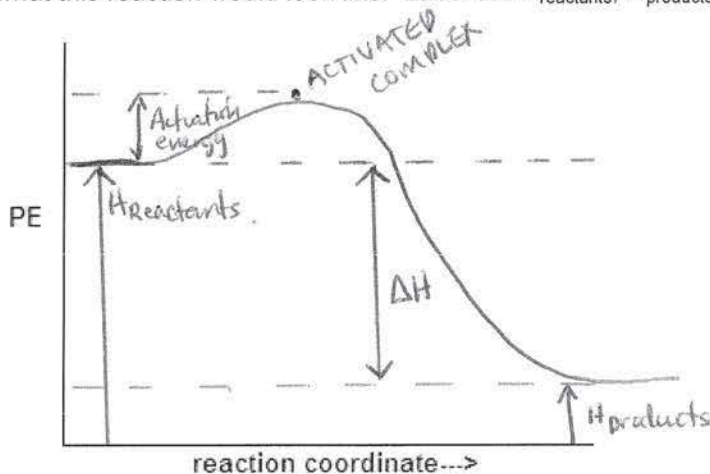
1) List one way in which the forward reaction can be made faster: INCREASE PRESSURE

2) Is this reaction exothermic or endothermic? Explain how you can tell. EXOTHERMIC Because Heat is one of the products

3) If this reaction were carried out in a calorimeter, would the temperature of the water in the calorimeter increase or decrease? Explain.

The temperature would increase b/c the water would absorb the heat.

4) Draw a PE Diagram sketch of what this reaction would look like. Label the  $H_{\text{reactants}}$ ,  $H_{\text{products}}$ ,  $H_{\text{activated complex}}$ ,  $\Delta H$  and activation energy.



5) List three ways in which this equilibrium can be stressed that will result in an increase in the  $\text{NH}_3(\text{g})$  concentration:

PRESSURE, Increase Reactants and Catalyst

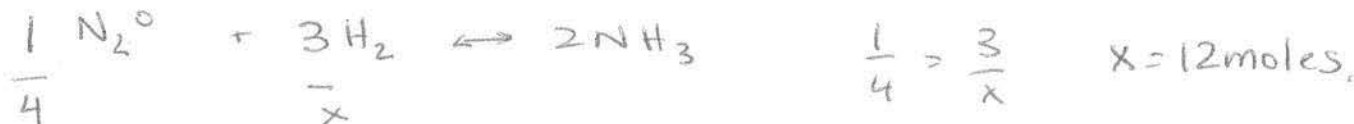
6) What kind of reaction is this? SYNTHESIS

7) Determine the charge of each species in the above reaction and write the:

a) oxidation half-reaction:  $\text{N}_2^0 \rightarrow 2\text{N}^{+3} + 6\text{e}^-$  reduction half-reaction:  $3\text{H}_2^0 + 6\text{e}^- \rightarrow 2\text{H}_3^{-1}$

b) oxidizing agent:  $\text{H}_2^0$  d) reducing agent:  $\text{N}_2^0$

9) How many moles of  $\text{H}_2(\text{g})$  are required to completely react with 4.0 moles of  $\text{N}_2(\text{g})$ ? Show your work.



10) How many moles of  $\text{NH}_3(\text{g})$  are formed when 3.0 moles of  $\text{H}_2(\text{g})$  are completely reacted with  $\text{N}_2(\text{g})$ ? Show work:



## Regents Exam In Chemistry Review Homework #4

Name \_\_\_\_\_

A solution contains 20. grams of  $\text{KNO}_3$  dissolved into 100. grams of water at  $40^\circ\text{C}$ .

- G 1) Is this solution saturated, unsaturated or supersaturated? Supersaturated
- 2) Explain how you can tell. The point is above the saturation point
- 3) By how many degrees does the solution have to be raised/lowered to make it saturated?  $5^\circ\text{C}$
- 4) How many grams can be added/ will precipitate to make the solution saturated? 5 grams will precipitate
- 5) Is the solution a homogenous or heterogeneous mixture? homogenous
- 6) Explain your answer to 5). it is the same throughout.
- 7) How can  $\text{KNO}_3$  be made more soluble in water? INCREASE TEMPERATURE
- 8) What is the name of the compound  $\text{KNO}_3$ ? Potassium nitrate
- 9) What is  $\text{KNO}_3$ 's formula mass? K 39 + N 14.00 + O (3x16) = 48 = 101g/mol
- 10) Is  $\text{KNO}_3$  an empirical or molecular formula? EMPIRICAL (SIMPLEST RATIO)
- 11) Explain your answer to 10). SIMPLEST RATIO (CANT BE REDUCED)
- 12) Determine the percent composition, by mass, of each element in  $\text{KNO}_3$ , showing all work:

$$\%K: \%K = \frac{\text{mass K}}{\text{mass KNO}_3} \times 100$$

$$= \frac{39.00}{101.00} \times 100 = 38.6\%$$

$$\%N: = \frac{\text{mass N}}{\text{mass KNO}_3} \times 100$$

$$= \frac{14.00}{101.00} \times 100 = 13.86\%$$

$$\%O: = \frac{\text{mass (3xO)}}{\text{mass KNO}_3} \times 100$$

$$= \frac{48.00}{101.00} \times 100 = 47.5\%$$

- 13) 5.0 moles of  $\text{KNO}_3$  are dissolved into 3.0 L of solution. Calculate the molarity of the solution, showing work:

$$\text{Molarity} = \frac{\text{moles}}{\text{LITERS}} = \frac{5.0 \text{ moles}}{3.0 \text{ L}} = 1.67 \frac{\text{moles}}{\text{LITERS}} \text{ or } 1.67 \text{ M}$$

- 14) How many grams of  $\text{KNO}_3$  are needed to make 2.0 L of 0.50 M  $\text{KNO}_3$  solution? Show all work:

$$\text{Molarity} = \frac{\text{moles}}{\text{LITERS}}$$

$$0.50 \text{ M} = \frac{x}{2.0 \text{ L}}$$

$$x = (2.0 \text{ L})(0.5 \frac{\text{moles}}{\text{L}})$$

$$x = 1 \text{ mole}$$

$$1 \text{ mole} = \frac{\text{given}}{\text{gfm}}$$

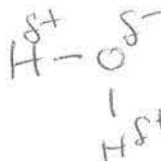
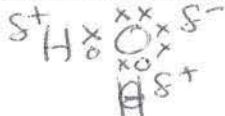
$$1 \text{ mole} = \frac{101.00 \text{ g}}{101.00} = \boxed{x = 101.00 \text{ g}}$$

## Regents Exam In Chemistry Review Homework #5

Name \_\_\_\_\_

### A) 100.0 grams of liquid water are at 0°C.

- 1) If heat is removed from this water, what phase change will occur? liquid  $\rightarrow$  solid
- 2) How many joules per gram are required to undergo this phase change? 334 J/g
- 3) How many joules are required for 100.0 g of water to undergo this phase change?  $q = m \Delta H_f (100)(334)$
- 4) What happens to the temperature of the water as it undergoes this phase change? STAYS THE SAME
- 5) Oxygen undergoes this phase change at 55 K. Convert this temperature to °C:  $K = ^\circ C + 273$   $55 = ^\circ C + 273$
- 6) Which molecule has stronger attractive forces, H<sub>2</sub>O or O<sub>2</sub>? H<sub>2</sub>O  $^\circ C = 55 - 273$   
 $= -218 K$
- 7) Draw the structural formula for a molecule of water:



- 8) Draw the dot diagram for a molecule of water:



- 9) Is a water molecule polar or nonpolar? Explain how you determined this.

polar, because it is a bent molecule with a lone pair of e<sup>-</sup>

- 10) What is the name for the type of attractive forces holding molecules of water together? Hydrogen bonding

- 11) What type of bond holds an H atom to an O atom? polar

- 12) How did you determine your answer to 11)? LEN

- 13) When NaCl dissolves in water, which ion is the oxygen end of the water molecule attracted to? Na<sup>+</sup>

- 14) Explain your answer to 13) Sodium is positively charged

- 15) When NaCl is dissolved into water, what happens to the freezing point of the water? decreases

k:

## Regents Exam In Chemistry Review Homework #6

Name \_\_\_\_\_

**200.0 grams of liquid water are heated from 20.0°C to 70.0°C.**

- 1) Is this a physical or chemical change? physical
- 2) Explain your answer to 1) IT IS STILL WATER
- 3) What happens to the viscosity of the water as it is being heated? Decreases
- 4) What happens to the vapor pressure of the water as it is being heated? Increases
- 5) What happens to the kinetic energy of the water as it is being heated? Increases
- 6) What happens to the entropy of the water as it is being heated? Increases
- 7) How many joules of energy must be added to the water to make it undergo this temperature change? Show work:

$$q = mc\Delta T$$

$$q = (200g) \left( \frac{4.185}{g^{\circ}C} \right) (70.0 - 20.0) = \text{joules}$$

- 8) Which will react faster? Na (s) + H<sub>2</sub>O (l) at 20°C, or Na (s) + H<sub>2</sub>O (l) at 70°C? \_\_\_\_\_
- 9) Explain your answer to 8) in terms of collision theory; The higher the average kinetic energy, the quicker they move.
- 10) 2 Na (s) + H<sub>2</sub>O (l) → H<sub>2</sub> + Na<sub>2</sub>O Complete the reaction.
- 11) What type of chemical reaction is this? SINGLE REPLACEMENT, REDOX
- 12) Identify the species being oxidized in question 10). Na<sup>0</sup>
- 12) Based on the Alternate Theory definition of acids and bases, is the water acting as an acid or base in the reaction in question 10? Explain.

Acid, because it's a proton donor. (gives H<sup>+</sup> to Na<sub>2</sub>O)

- 13) Is the product you wrote the formula for in question 10 an acid or a base? Explain how you can tell:

Base, because Na<sub>2</sub>O accepts a H<sup>+</sup> from H<sub>2</sub>O

- 14) Water has a pH of 7. If the concentration of OH<sup>-</sup> ions increases 1000 times as Na is added to the water, what will the new pH be?

pH=10

- 15) What color will methyl orange be in this pH? yellow.

- 16) If 20.0 mL of 4.0 M H<sub>2</sub>SO<sub>4</sub> are needed to completely neutralize 80.0 mL of NaOH solution, then what is the molarity of the NaOH solution? Show your work.

$$\begin{aligned} \#H^+ M_A V_A &= M_B V_B \#OH^- \\ \frac{(2)(4.0)(20.0)}{(80.0)(1)} &= M_B \\ M_B &= 2M \end{aligned}$$

## Regents Exam In Chemistry Review Homework #7

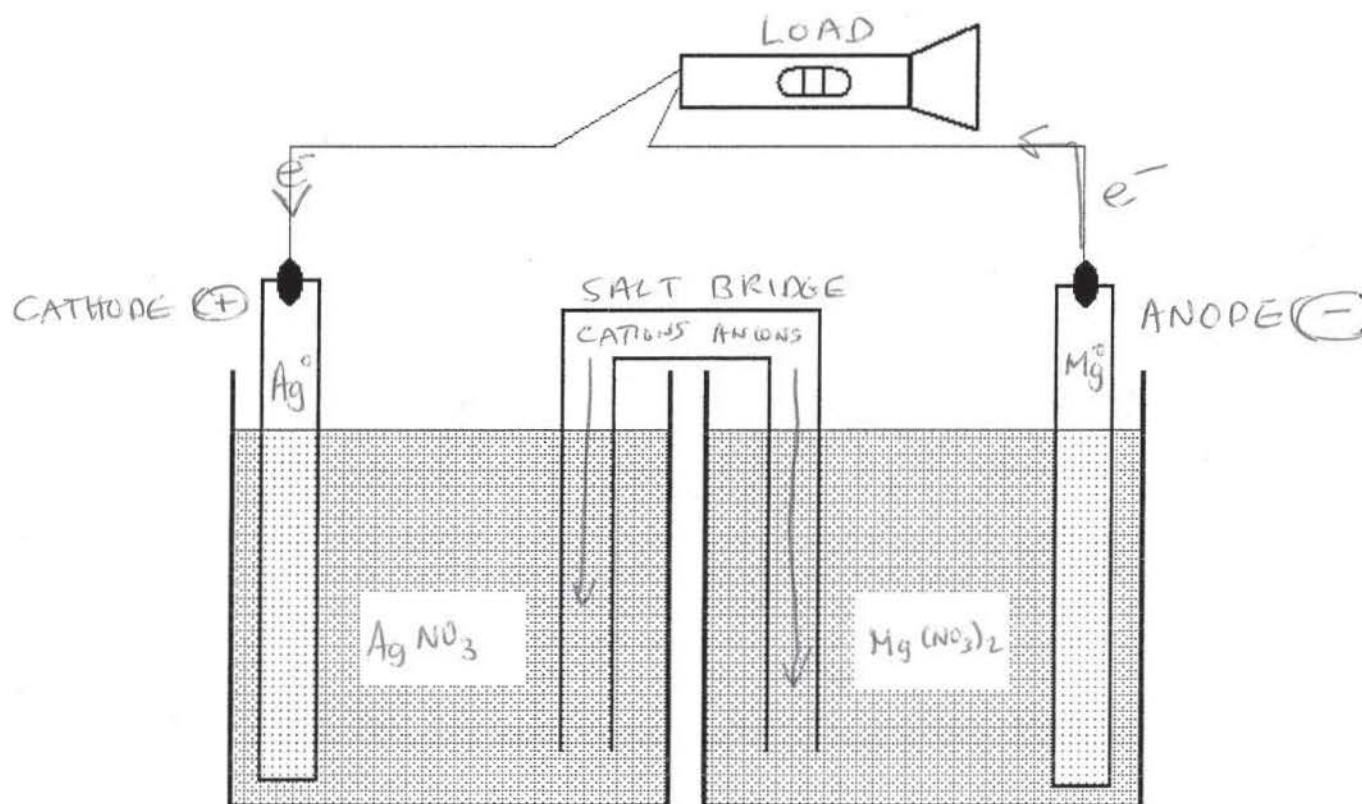
Name \_\_\_\_\_

A battery is made using the reaction  $\text{Mg} + \text{AgNO}_3 \rightarrow \text{Mg}(\text{NO}_3)_2 + \text{Ag}$ .

- 1) Select a metal on Table J that would also work in this reaction: Cu or above
- 2) Explain why the metal you chose would work: because it is above Ag
- 3) Balance the reaction 1  $\text{Mg}^0 +$  2  $\text{Ag}^{+1}\text{NO}_3 \rightarrow$  1  $\text{Mg}^{+2}(\text{NO}_3)_2 +$  2  $\text{Ag}^0$
- 4) Write the charge of each species in this reaction:  

$$\begin{array}{ccccccc} 0 & +1 & -1 & +2 & -1 & 0 \\ \text{Mg} & + \text{AgNO}_3 & \rightarrow & \text{Mg}(\text{NO}_3)_2 & + \text{Ag} \end{array}$$
- 5) Write the oxidation half-reaction:  $\text{Mg}^0 \rightarrow \text{Mg}^{+2} + 2e^-$
- 6) Write the reduction half-reaction:  $\text{Ag}^{+1} + 1e^- \rightarrow \text{Ag}^0$
- 7) Identify the oxidizing agent:  $\text{Ag}^{+1}$  Reducing Agent:  $\text{Mg}^0$  Spectator Ion:  $\text{NO}_3^{-1}$
- 8) Draw and label a voltaic cell based on this reaction. Label the following:

Anode, cathode, + electrode, - electrode, direction electrons take, composition of all electrodes and solutions, load, salt bridge, direction that anions go across the salt bridge and direction that cations go across the salt bridge.



# Regents Exam In Chemistry Review Homework #8

Name \_\_\_\_\_

For the reaction  $\text{KCl} \rightarrow \text{K} + \text{Cl}_2$ :

- 1) Balance the reaction:  $2 \overset{+1}{\text{K}}\overset{-1}{\text{Cl}} \rightarrow 2 \overset{0}{\text{K}} + 1 \overset{0}{\text{Cl}_2}$
- 2) Identify what type of reaction is represented here: Decomposition
- 3) What phase does the KCl have to be in in order to electrolytically decompose the compound? l
- 4) K will form at the - charged electrode (the cathode), where reduction occurs.
- 5)  $\text{Cl}_2$  will form at the + charged electrode (the anode), where OXIDATION occurs.
- 6) Write the oxidation half-reaction:  $2\text{Cl}^- \rightarrow \text{Cl}_2^0 + 2e^-$
- 7) Write the reduction half-reaction:  $2\text{K}^+ + 2e^- \rightarrow 2\text{K}^0$
- 8) How many moles of  $\text{Cl}_2$  will form if 4.0 moles of KCl are decomposed? Show your work.

$$\frac{2^0}{4^0} = \frac{1^0}{x} \quad \frac{2}{4} = \frac{1}{x} = 2x = 4 \quad x = 2 \text{ moles}$$

- 9) Sargent-Welch has ordered 100. moles of K from your company. How many moles of KCl must be decomposed to make the order? Show your work.

$$\begin{array}{l} 2 : 2 \\ 1 : 1 \end{array} \quad \begin{array}{l} 1 : 1 \\ 100 : x \end{array} \quad \frac{1}{100} = \frac{1}{x} \quad x = 100 \text{ moles}$$

- 10)  $\text{Cl}_2$  is a gas. It can be collected by trapping it under water. Will the  $\text{Cl}_2$  be soluble in the water? Explain, in terms of molecular polarity.

NO, because  $\text{Cl}_2$  is non-polar. Like dissolves like

- 11) What is the name of the group on the Periodic Table that K belongs to? Alkali

- 12) Write the dot diagram for an atom of K:  $\text{K}^0$

- 13) Write the dot diagram for a molecule of  $\text{Cl}_2$ :  $\overset{\times \times}{\underset{\times \times}{\text{Cl}}} : \overset{\times \times}{\underset{\times \times}{\text{Cl}}} :$

- 14) Is this reaction a physical or chemical change?

chemical, because there is a change in the identity

- 15) Why is this reaction considered a redox reaction?

Because one is losing and one is gaining  
electrons